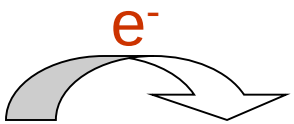
Oxidation & Reduction

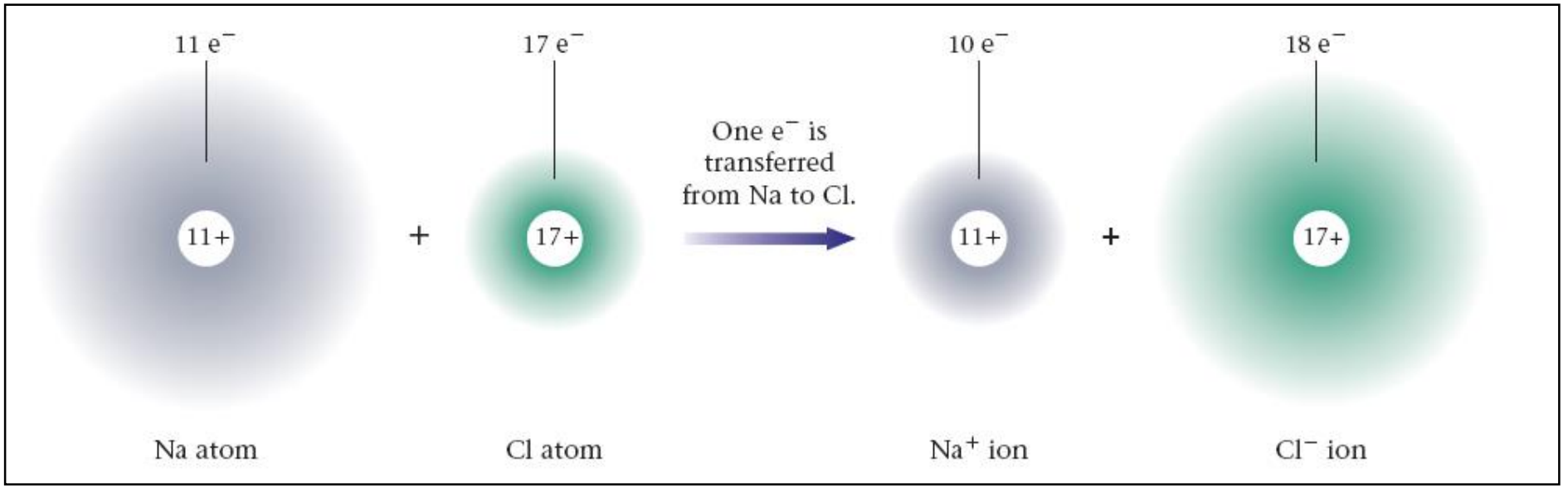
Oxidation and Reduction reactions (redox)



$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

$$Na \rightarrow Na^+ + e^-$$

$$Cl + e^- \rightarrow Cl^-$$



Oxidation and Reduction reactions (redox)

oxidation: it is the loss of electrons.

$$Na \rightarrow Na^{+} + e^{-}$$

reduction: it is the gain of electrons.

$$Cl + e^{-} \rightarrow Cl^{-}$$

Remember – **LEO** says **GER**.
Loss of **E**lectrons is **O**xidation
Gain of **E**lectrons is **R**eduction.

Oxidation and Reduction reactions (redox)

Metal + Nonmetal : Transfer of electrons

Oxidation and Reduction reactions (redox)

Oxidation and reduction *always* occur together.

(The lost e^- must go somewhere!)

Oxidation and Reduction reactions (redox)

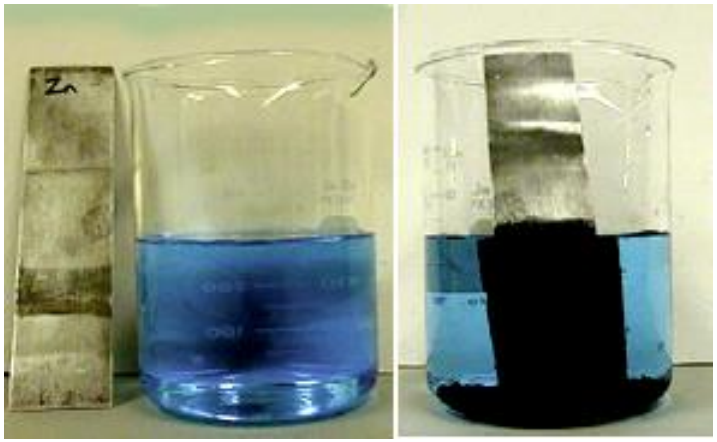
oxidation: it is the loss of electrons.

reduction: it is the gain of electrons.

$Zn(s) + Cu^{2+}(aq) \rightarrow Zn^{2+}(aq) + Cu(s)$ redox reaction

$Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$ Zn is oxidized **(reducing agent)**

$Cu^{2+}(aq) + 2e^- \rightarrow Cu(s)$ Cu^{2+} is reduced **(oxidizing agent)**

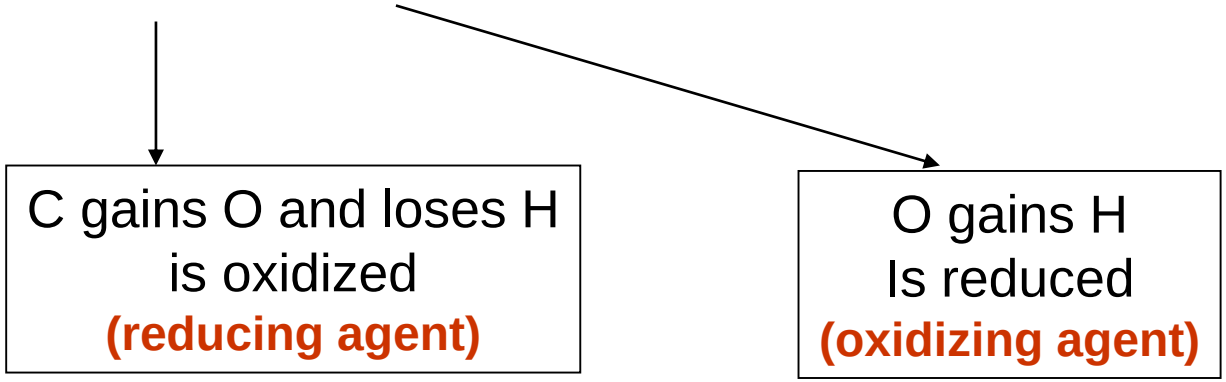


Oxidation and Reduction reactions (redox)

oxidation: is the gain of oxygen / loss of hydrogen.

reduction: is the loss of oxygen / gain of hydrogen.

$CH_4(s) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ redox reaction



single replacement reaction and combustion reactions $\rightarrow$ redox reactions

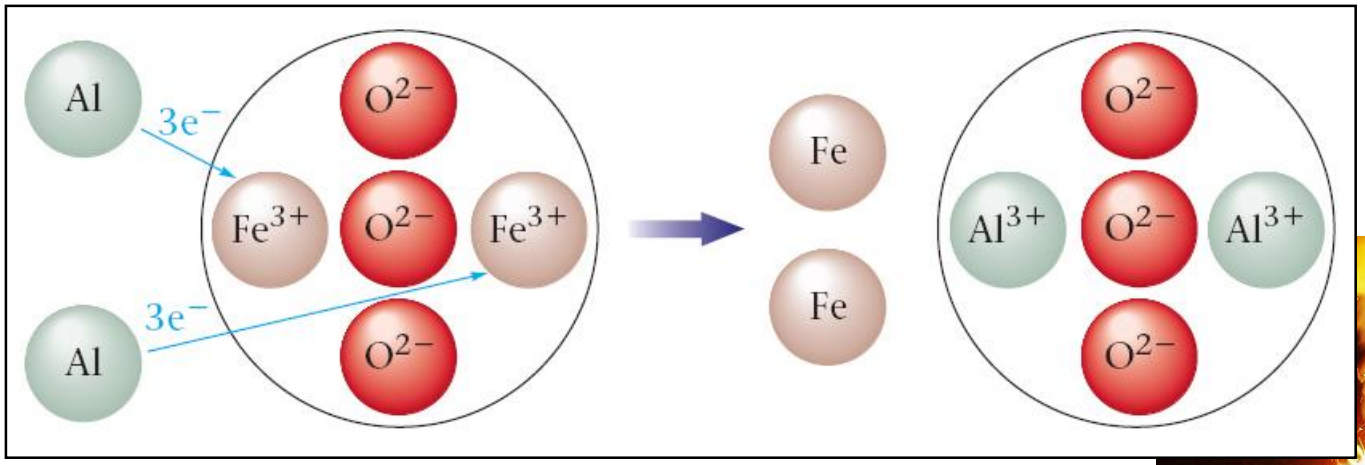
double replacement reactions $\rightarrow$ non redox

Oxidation and Reduction reactions (redox)

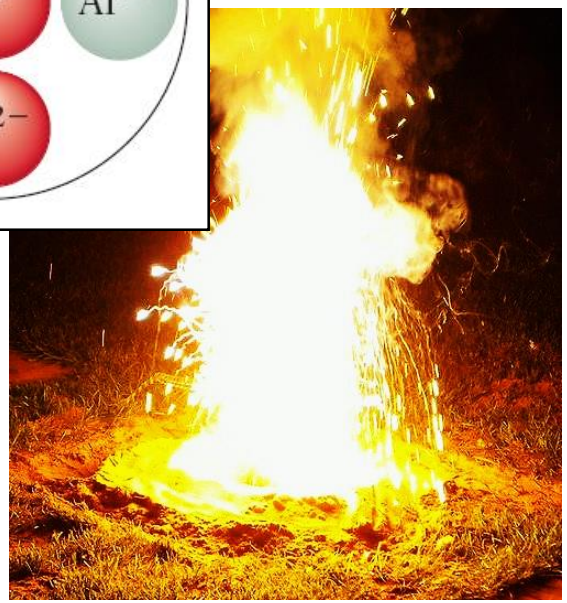
Example 2:

- $2\ Al(s) + Fe_2O_3(s) \rightarrow 2\ Fe(s) + Al_2O_3(s)$

  is oxidized is reduced



- Called the *Thermite reaction*.
- Let's just say it's vigorous!



Oxidation and Reduction reactions (redox)

Example 3:

$$Cu(s) + 2\ Ag^{+}(aq) \rightarrow 2\ Ag(s) + Cu^{2+}(aq)$$

is oxidized is reduced

Oxidation and Reduction reactions (redox)

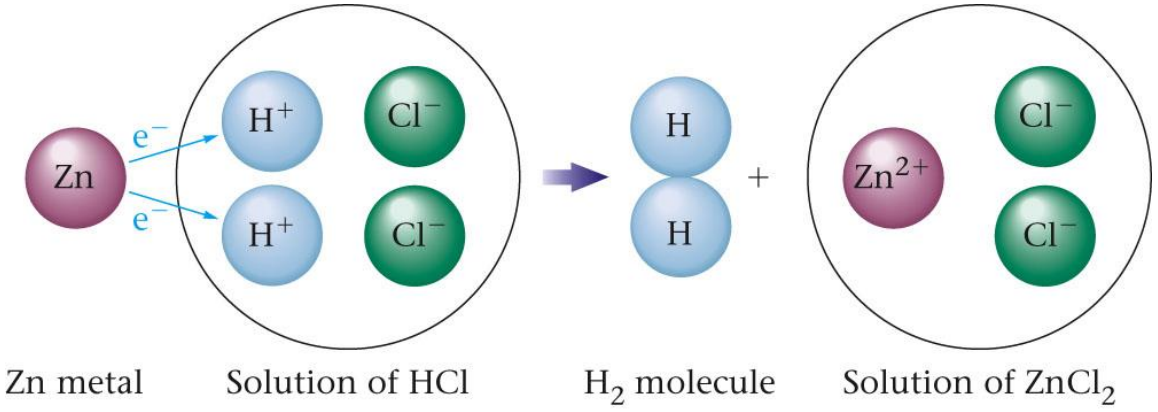
Example 4:

$$Zn(s) + 2\ HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$$

$$Zn(s) + 2\ H^+(aq) + \cancel{2Cl^-(aq)} \rightarrow H_2(g) + Zn^{2+}(aq) + \cancel{2Cl^-(aq)}$$

$$Zn(s) + 2\ H^+(aq) \rightarrow H_2(g) + Zn^{2+}(aq)$$

is oxidized is reduced



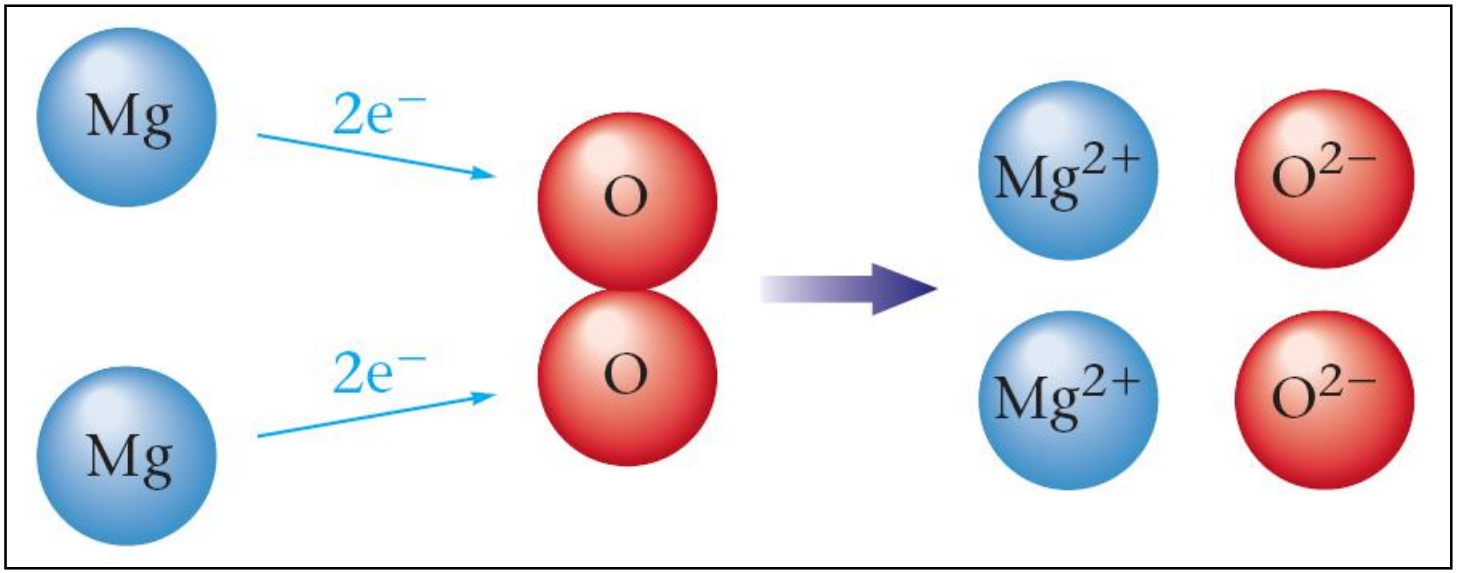
Note: this reaction also shows the fourth driving force of a reaction, namely, *the formation of a gas*.

Oxidation and Reduction reactions (redox)

Example 5:

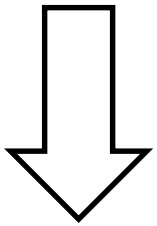
$$2\ Mg(s) + O_2(g) \rightarrow 2\ MgO(s)$$

is oxidized is reduced



Oxidation States (Oxidation numbers)

Assigning charges to the various atoms in a compound.



Keep track of electrons in redox reactions.

Rules for assigning oxidation states

1. Charge (oxidation state) of a uncombined element is zero.

H_2, Cl_2, Ar, Na, K

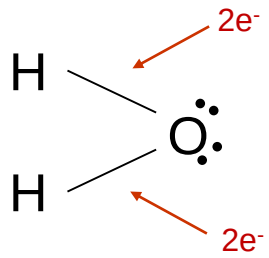
2. The oxidation state of a monatomic ion is the same as its charge.

NaCl Na^+ Cl^- ⟶ Na: +1 Cl: -1

Al_2S_3 Al^{3+} S^{2-} ⟶ Al: +3 S: -2

Rules for assigning oxidation states

For covalent compounds assume the most electronegative atom controls or possesses the shared electrons.



O gained two e^- from H → Oxidation state = -2

H lost one e^- → Oxidation state = +1

4. The oxidation state of H is +1 and O is -2 in covalent compounds.

Exception: Peroxide (O_2^{2-}) = -1 H_2O_2

Rules for assigning oxidation states

The most electronegative elements: F, O, N, and Cl

F: -1, O: -2, N: -3, Cl: -1

5. If two of these elements are found in the same compound, we assign them in order of electronegativity.

F > O > N > Cl

NO_2

O: 2 × (-2) = -4 So N must be +4

Rules for assigning oxidation states

6. Sum of oxidation states = 0 in a neutral compound.

7. Sum of oxidation states = charge in an ion.

NO_3^- O: 3 × (-2) = -6

N must be +5 for an overall charge of -1 for NO_3^-.

SO_4^{2-} O: 4 × (-2) = -8

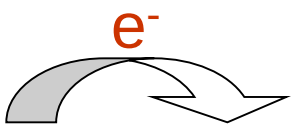
S must be +6 for an overall charge of -2 for SO_4^{2-}.

Rules for assigning oxidation states

$K_2Cr_2O_7$	K = +1; Cr = +6; O = –2
CO_3^{2-}	C = +4; O = –2
MnO_2	Mn = +4; O = –2
PCl_5	P = +5; Cl = –1
SF_4	S = +4; F = –1

Oxidation-Reduction Reactions

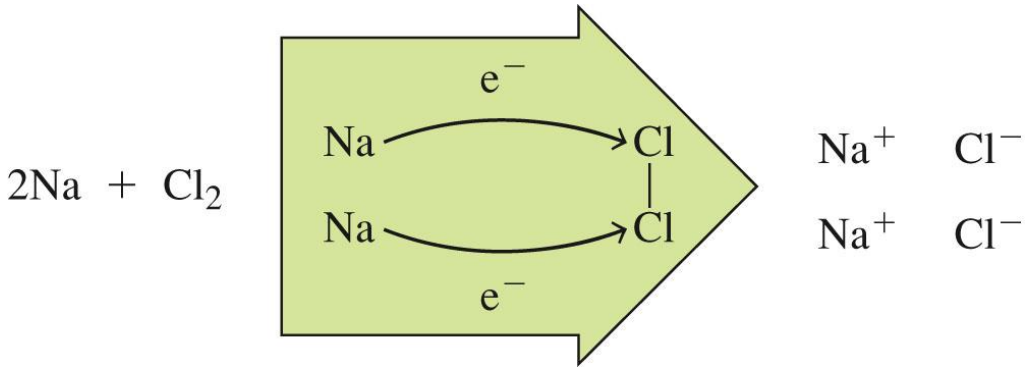
In some redox reactions ions are produced.



$$2Na(s) + Cl_2(g) \longrightarrow 2NaCl(s)$$

Oxidation state: 0 0 +1 -1

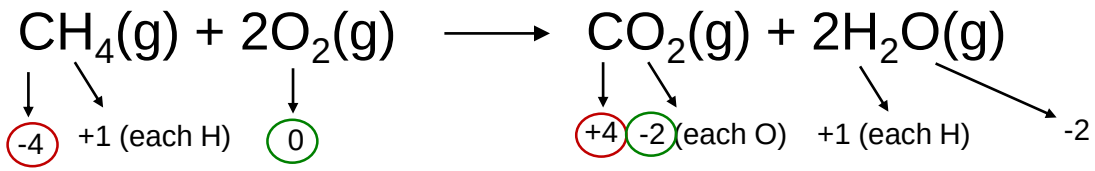
Element Element Na^+ Cl^-

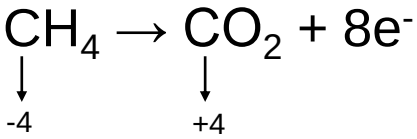


Oxidation-Reduction Reactions

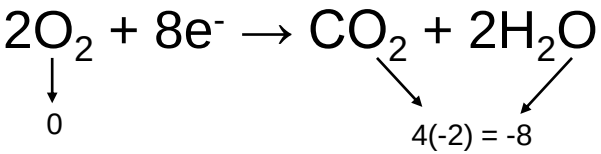
In some redox reactions ions are not produced (all nonmetals).

Oxidation state:





C is oxidized.
CH_4 is a reducing agent.



O is reduced.
O_2 is an oxidizing agent.

Oxidation-Reduction Reactions

Oxidation: is an increase in oxidation state (a loss of e^-).

Reduction: is a decrease in oxidation state (a gain of e^-).

Oxidizing agent (electron acceptor): the reactant containing the element that is reduced.

Reducing agent (electron donor): the reactant containing the element that is oxidized.

$CH_4 \rightarrow CO_2 + 8e^-$

C is oxidized.
CH_4 is a reducing agent.

$2O_2 + 8e^- \rightarrow CO_2 + 2H_2O$

O is reduced.
O_2 is an oxidizing agent.

Half-Reaction Method for balancing

$$Ce^{4+}(aq) + Sn^{2+}(aq) \rightarrow Ce^{3+}(aq) + Sn^{4+}(aq)$$

$Ce^{4+}(aq) + e^- \rightarrow Ce^{3+}(aq)$ Reduction half-reaction

$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$ Oxidation half-reaction

Number of Electrons lost ⟷ must be equal to ⟷ Number of Electrons gained

Multiply by 2: $2Ce^{4+}(aq) + 2e^- \rightarrow 2Ce^{3+}(aq)$

$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq) + 2e^-$

$2Ce^{4+}(aq) + Sn^{2+}(aq) + \cancel{2e^-} \rightarrow 2Ce^{3+}(aq) + Sn^{4+}(aq) + \cancel{2e^-}$

$2Ce^{4+}(aq) + Sn^{2+}(aq) \rightarrow 2Ce^{3+}(aq) + Sn^{4+}(aq)$

Half-Reaction Method for balancing

1. Identify and write the equations for the oxidation and reduction half–reactions.

2. For each half–reaction:
   A. Balance all the elements except H and O.
   B. Balance O using H_2O.
   C. Balance H using H^+.
   D. Balance the charge using electrons.

3. If necessary, multiply one or both balanced half–reactions by an integer to equalize the number of electrons transferred in the two half–reactions.

4. Add the half–reactions, and cancel identical species.

5. Check that the elements and charges are balanced.

Half-Reaction Method for balancing

$$Cr_2O_7^{2-}(aq) + SO_3^{2-}(aq) \rightarrow Cr^{3+}(aq) + SO_4^{2-}(aq)$$

How can we balance this equation?

1. Separate into half-reactions.

2. Balance elements except H and O.

$$Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq)$$

$$SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq)$$

Half-Reaction Method for balancing

3. Balance O's with H_2O and H's with H^+.

$$14H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(aq)$$

$$H_2O(aq) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2H^+(aq)$$

4. How many electrons are involved in each half-reaction? Balance the charges.

$$6e^- + 14H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(aq)$$

$$H_2O(aq) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-$$

Half-Reaction Method for balancing

5. Multiply whole reactions by a whole number to make the number of electrons gained equal the number of electrons lost.

$$6e^- + 14H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + 7H_2O(aq)$$

$$3 \times (H_2O(aq) + SO_3^{2-}(aq) \rightarrow SO_4^{2-}(aq) + 2H^+(aq) + 2e^-)$$

6. Combine half-reactions cancelling out those reactants and products that are the same on both sides, especially the electrons.

$$\overset{8}{6e^-} + 14H^+(aq) + Cr_2O_7^{2-}(aq) \rightarrow 2Cr^{3+}(aq) + \overset{4}{7}H_2O(aq)$$

$$3H_2O(aq) + 3SO_3^{2-}(aq) \rightarrow 3SO_4^{2-}(aq) + 6H^+(aq) + 6e^-$$

$$Cr_2O_7^{2-} + 3SO_3^{2-} + 8H^+ \rightarrow 2Cr^{3+} + 3SO_4^{2-} + 4H_2O$$